Student Textbook page 130

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1. 586 mL

2. 92 mL

3. 14 mL

4. 1.2 atm

5. 84 °C

6. 56.9 °C

Student Textbook page 133

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7. 250 mL

8. 5.0 mL

9. 700 mL

Student Textbook page 133

Q1. The combined gas law combines Boyle's law and Charles's law.

Q2. The law of combining volumes applies to chemical reactions involving gases.

Student Textbook page 135–136

Q3. Avogadro's law states that equal volumes of all ideal gases at the same temperature and pressure contain the same number of molecules. That does not mean that they have the same mass. Scientists had to separate the concepts of ratios of volumes and ratios of mass. They are not directly related, only related through molar mass.

Q4. According to Avogadro's law, equal volumes of gases as the same temperature and pressure will contain the same number of molecules. Therefore, the volume of nitrogen at that temperature and pressure will also be 29 L.

Q5. Since gas volumes are affected by temperature and pressure, it allows scientists to communicate information about gas volumes at these convenient temperatures.

Student Textbook page 138

1. As the gas is heated, it expands and some of the gas particles leave the container.
2. She probably heated the can too long. The water boiled away, so the can was full of air when she transferred it to the cold water. A vacuum does not form, and the can does not crush. (It would, however, compress somewhat as the air in the can cools.)
3. Hydrogen : chlorine : hydrogen chloride = 1 : 1 : 2. This comes from the balanced chemical reaction.
4. 6.98 g / 70.90 g/mol = 0.0984 mol

   22.4 L/mol x 0.0984 mol = 2.21 L
5. The ratio of hydrogen gas water vapour is 1 to 1. You make 2.3 L of water vapour.
6. P = 60 mL × 100 kPa × 328.15 K /(120 mL × 298.15 K) = 55 kPa
7. This describes SATP so the answer should be 24.8 L/mol. Calculations with the data give:

   V = 2.27 L × 101.325 kPa × 298.15 K /(100 kPa × 273.15 K) = 2.51 L

   The amount of gas is 6.98 g / 35.45 g/mol = 0.197 mol

   Molar volume is 2.51 L / 0.197 mol = 25.5 L/mol
8. T = 110 kPa × 24 mL × 273 K / (101.3 kPa × 48 mL) = 1.5×10^2 K (-1.3×10^2 °C)
9. V = 100.8 kPa × 4.2 L × 248 K / (271 K × 103 kPa) = 3.8 L

10. If you are looking for the relationship between molar volume and pressure then you have to control the temperature. The manipulated variable is the pressure and the responding variable is the molar volume. Conversely, if you are looking for the relationship between molar volume and temperature, then control the pressure. The manipulated variable is the temperature and the responding variable is the molar volume. This is a good exercise for a spreadsheet solution. Reading across the rows works for the pressure-molar volume relationship. Reading down the columns works for the temperature-molar volume ratio. Charting these numbers as a bar chart or line graph would work just fine. If the students had coloured diagrams similar to Figures 4.2, 4.3, and 4.4, that would give them what they need to do the presentation.

11. As a balloon rises in the atmosphere, it will be subjected to a decrease in pressure (atmospheric) and a decrease in temperature. The decreased pressure will lead to an increase in the balloon's volume since pressure and volume are inversely proportional. The decrease in temperature will lead to a decrease in volume since volume and temperature are directly proportional. Whether the balloon expands (and perhaps eventually pops) or gets smaller will depend on whether the pressure decreases at a greater or lesser rate than the Kelvin temperature decreases. If the temperature decreases at a faster rate, then the balloon will shrink in size; if the pressure decreases at a faster rate, then the balloon's volume will increase. In reality, balloons will increase in size and eventually pop. Weather and other instrument carrying balloons are designed to release some of the trapped gases when the difference in pressure becomes too great instead of popping violently.

Student Textbook page 141

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10. 118 kPa

11. 1.4×10^2 L

12. 1.01×10^{-3} mol

13. 2.8 L

14. 3.0×10^5 K

15. 166 °C

16. 38.0 g/mol

Answers to Questions for Comprehension

Student Textbook page 141

Q6. To solve for the universal gas constant, start with molar volume at STP.

$n = 1.00$ mol

$P = 101.325$ kPa

$T = 273.15$ K

$V = 22.4$ L

$R = ?$

$PV = nRT$

$$R = \frac{PV}{nT}$$

$$= \frac{(101.325 \text{ kPa})(22.4 \text{ L})}{(1.00 \text{ mol})(273.15 \text{ K})}$$

$$= 8.31 \frac{\text{kPa} \cdot \text{L}}{\text{mol} \cdot \text{K}}$$

Q7. If you use 8.314 kPa•L/mol•K as the universal gas constant, then you must use matching units so they will cancel.

Student Textbook page 142

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17. 1.78 g/L

18. 0.17 g/L

19. 97.6 kPa

Student Textbook page 144

Q8. If you ignore all other gases in the atmosphere, the partial pressure of oxygen is the difference between the total pressure (102 kPa) and the pressure of nitrogen gas (71 kPa) or 102 kPa – 71 kPa = 31 kPa. On a humid day, if the total pressure remained at 102 kPa, the partial pressures of nitrogen and oxygen gases would be lower because the total pressure would include the partial pressure exerted by water vapour.

Q9. Two factors that must be considered when collecting gas using water displacement are: the possibility that some of the gas will dissolve in the water and the possibility that the pressure of the gas collected is affected by the partial pressure of water vapour.

Q10. The pressure of the "dry" hydrogen is the total pressure of the gas collected minus the vapour pressure of water at that temperature.

P_{H_2} = 99.8 kPa – 2.33 kPa = 97.5 kPa.

Student Textbook page 150

1. The ideal gas law uses moles of a gas not the mass of the gas. This makes the law work no matter how heavy a particle is.

2. One method would be to calculate the molar volume of a gas at particular conditions of temperature, pressure, and volume, and then to take a sample of a gas at those same conditions. Measure the temperature, pressure, and volume, and use the ideal gas law to calculate the number of moles. Measure the mass of the gas as a separate step. The molar mass is the mass divided by the amount of the gas.

   Another method is to produce a gas from a chemical reaction and use volume ratios to determine the amount of gas produced. Measure the mass of the gas, and again you have the numbers needed to find the molar mass of the gas.

3. No. Hydrogen sulfide, like all acid forming substances, is highly soluble in water. Only gases that do not dissolve in water can be collected with this method.

4. 22.4 L + 22.4 L + 33.6 L + 22.4 L = 100.8 L and for significant figures the answer is 1.0×10^2 L

5. Using the chart, the partial pressure of water at this temperature is 0.16 kPa. The pressure of the gas is therefore 100.2 kPa – 2.41 kPa = 97.8 kPa.

6. The diagrams suggest that there is lots of space between gas particles, even after adding the two gases together so the total pressure is the sum of the partial pressures.

7. $5.25 \times 10^{27} / 6.02 \times 10^{23} = 8721$ mol $\times$ 24.8 L/mol = 2.16×10^5 L

8. Solving the ideal gas law for n gives n = 0.0135 mol. So M = 0.087 g / 0.0135 mol = 64 g/mol.

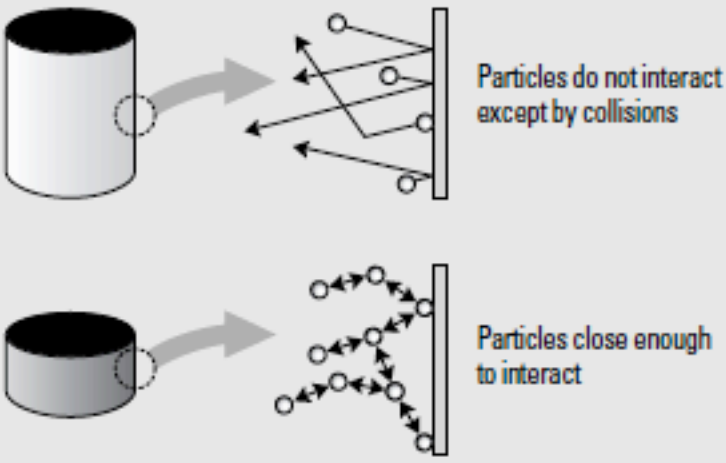
9. The volume inside the cylinder is too high. When collecting a gas using the downward displacement of water, you must equalize the internal and external pressure to obtain accurate results for the quantity of gas produced. The cylinder should be carefully lifted until the level of water outside the cylinder is equal to the water level inside the cylinder. When the water levels are equal, the pressure in the cylinder is the same as atmospheric pressure, which can be easily measured. Currently, the pressure inside the cylinder is higher than the atmospheric pressure outside the cylinder, and therefore the volume of gas will be less than it would be at atmospheric pressure (since pressure and volume are inversely proportional). When they calculate n, their value will be too low.

10. $V = \frac{nRT}{P} = \frac{(1.2)(8.314)(298.15)}{100} = 30 \text{ L}$

Student Textbook pages 152–153

1. The combined gas law is more general, (less constrained). You do not need a constant pressure, as in Charles's law, or a constant temperature, as in Boyle's law.
2. In chemical reactions, volumes add in the same proportions as the mole ratios. This means that equal amounts (moles) of any gas occupy equal volumes.
3. Since equal amounts of any gas have the same volume, we expect the volume to be related to the amount of gas. The diagram shows the amount increasing and the volume responding accordingly.
4. Having the same molar volume means having the same number of particles in a given volume. Density is mass divided by volume and different densities result because different gas particles have different masses.

5. This assures that the mole ratios are accurate.
6. The kinetic molecular theory describes an ideal gas as one where the particles can be thought of as point masses, there are no chemical interactions between the particles, and there is lots of space between the particles.
7. Since the space between gases is so large in comparison to the size of the particles there is plenty of room for different types of particles. Each type of particle exerts an average force on the container in proportion to the number of particles there are.
8. Every gas has the same molar volume at a given set conditions. STP and SATP are special conditions.
9. **(a)** 8.314 kPa • L/mol • K × 1 atm/101.325 kPa = 0.08205 atm • L/mol • K

   (b) 8.314 kPa • L/mol • K × 760 mmHg/101.325 kPa = 62.4 mmHg • L/mol • K
10. The two things that cause a gas to behave differently than an ideal gas are high pressure and low temperature. High-pressure squeezes the particles together so that the space between the atoms is not large in comparison to the size the particles and particle interactions begin to matter. At low temperatures the particles slow down and come closer together, again, causing particle interactions to become important.
11. 

The diagram shows the particles much closer together at high pressure than low pressure. The attractions between particles take something away from the pressure they might exert on the walls.

12. At high pressure the particles are close enough to attract one another, which reduces the force with which they hit the container's walls. As a result, the pressure is reduced. Calculations of n with the ideal gas law will be low as a result.

13. The volume is directly proportional to the temperature and inversely proportional to the pressure. The dropping temperature would tend to reduce the volume of the balloon, but the dropping pressure would tend to increase it. In absolute terms, the pressure changes are likely bigger than the temperature changes, so the balloon would probably increase in volume.

Answers to Applying Concepts Questions

14. When a liquid changes to gas, particles absorb enough energy to move away from the surface of the liquid and fly free. This does not change the temperature of the substance but in a closed container would drastically increase pressure because the volume has a tendency to increase dramatically. Adding heat can trigger a chemical reaction that creates gases. The heat produced by compressing the fuel is enough to ignite diesel fuel in a vehicle.

15. Avogadro's idea that equal volumes of gases contain the same number of particles explains why they combine in small (mole) ratios, which agrees with Dalton's ideas.

16. Assuming the containers are sealed:

(a) The volumes of the two gases are equal.

(b) The amounts of the two gases are equal.

(c) The mass of the oxygen gas will be eight times the mass of the helium gas.

(d) Assuming that the container has a constant volume and the amounts of each gas are still the same, the pressure of the oxygen gas would be less than the pressure of the helium gas, and since the temperature has decreased, the average kinetic energy of the oxygen particles is less than the average kinetic energy of the helium particles.

17. It means that the gas has no (or a very small amount) of water particles mixed in with it.

Answers to Problem Solving Questions

18. Use the combined gas law. $V_2 = P_1V_1T_2/P_2T_1$; V_2 = (1.33 atm × 101.325 kPa/atm × 532 mL × 273.15 K)/(101.325 kPa × (31.4 °C + 273.15) K = 635 mL

19. 21 mL

20. Use the combined gas law to find the total volume of the Helium. $V_2 = P_1V_1T_2/P_2T_1$; V_2 = (25 atm × 100 L × (25 °C + 273.15) K)/(1.05 atm × (20 °C + 273.15) K) = 2421.56 L. Now subtract the 100 L that are still in the cylinder. 2421.56 L – 100 L = 2321.56 L. Correct for significant figures. $V = 2.3 \times 10^3$ L.

21. Assuming an ideal gas (which may be a poor assumption here due to such a high pressure): Find the amount: n = 265 g / 32.00 g/mol = 8.28 mol
Then T = 20 atm × 101.325 kPa/atm × 10 L / (8.314 kPa • L/mol • K × 8.28 mol) = 294 K

22. Assuming an ideal gas (which may be a poor assumption here due to such a high pressure):

n = 200 bar × 100 kPa/bar × 5.4 L / (8.314 kPa • L/mol • K × 295.15 K) = 44.0 mol

Then m = 44.0 mol × 32.00 g/mol = 1408 g = 1.4 kg

23. The volume of the cylinder is $3.14 \times (11.5\text{cm})^2 \times 140\text{ cm} = 58137\text{ cm}^3$

This is 58.1 L

$n = 108\text{ kPa} \times 58.1\text{ L} / (8.314 \times 294\text{K}) = 2.57\text{ mol} \times 44.02\text{ g/mol} = 113\text{ g}$

24. Find the amount: $n = 1.25\text{ g} / 4.00\text{ g/mol} = 0.3125\text{ mol}$

Then: $22.4\text{ L/mol} \times 0.3125\text{ mol} = 7.00\text{ L}$

And $D = 1.25\text{ g} / 7.00\text{ L} = 0.179\text{ g/L}$

25. Use one mole: $V = 1\text{ mol} \times 8.314 \times 362\text{ K} / 155\text{ kPa} = 19.4\text{ L}$

$m = 1\text{ mol} \times 70.90\text{ g/mol} = 70.90\text{ g}$ and $D = 70.90\text{ g} / 19.4\text{ L} = 3.6\text{ g/L}$

26. (a) $C_3H_{8(g)} + 5O_{2(g)} \rightarrow 3CO_{2(g)} + 4H_2O_{(g)}$

(b) Using Avogadro's law: 1 L of propane will produce 3 + 4 = 7 L of products.

27. $n = 103\text{ kPa} \times 0.800\text{ L} / (8.314 \times 351\text{ K}) = 0.0282\text{ mol}$

$M = 2.366\text{ g} / 0.0282\text{ mol} = 83.4\text{ g/mol}$. The gas is krypton.

28. $T = 206\text{ kPa} \times 4.00\text{ L} / (2.17\text{ mol} \times 8.314) = 45.7\text{ K} = -227\text{ °C}$

29. Use the combined gas law and solve for V_2/V_1.

$V_2/V_1 = (P_1 T_2/P_2 T_1) = (P_1\ (75 + 273.15)\text{ K})/(2P_1\ (25 + 273.15)\text{ K}) = 0.58$

30. 4.0% of 0.75 atm = 0.030 atm or 3.04 kPa

31. Vapour pressure is 2.81 kPa = 21.1 mm Hg.

Dry pressure is 750 mmHg – 21.1 mmHg = 729 mmHg

New volume is: $V = 729\text{ mmHg} \times 55\text{ mL} \times 313\text{ K} / (296\text{ K} \times 775\text{ mmHg}) = 54.7\text{ mL}$

32. Use the combined gas law to find the final pressure (partial pressure) for the individual gases and then add the partial pressures to obtain the total pressure.

$P_2 = P_1 V_1 T_2 / V_2 T_1$;

$P_{2(Ar)}$ = (1.2 atm × 0.600 L (27 + 273.15) K)/ (0.400 L (277 + 273.15) K) = 0.9820 atm

$P_{2(O_2)}$ = (502 mmHg (1 atm/760 mmHg) 0.200 L (27 + 273.15) K)/(0.400 L (127 + 273.15) K) = 0.2477 atm

P_2(total) = 0.9820 atm + 0.2477 atm = 1.2297 atm

Corrected for significant figures: $P_{2(total)}$ = 1.2 atm or 1.2×10^2 kPa

33. Vapour pressure is 2.81 kPa.

Dry pressure is 738 mmHg / 760 mmHg × 101.3 kPa/atm = 98.37 kPa – 2.81 kPa = 95.56 kPa

n = 95.56 kPa × 0.523 L / (8.314 × 296 K) = 0.0203 mol

m = 0.0203 mol × 26.04 g/mol = 0.53 g

34. (a) Find the mass: 2.201 g – 2.150 g = 0.051 g

Find the amount: n = 102 kPa × 0.664 L / (8.314 kPa • L/mol • K × 323.15 K) = 0.0252 mol

The molar volume: = 0.664 L / 0.0252 mol = 26.3 L/mol

The molar mass: M = 0.051g / 0.0252 mol = 2.02 g/mol

(b) The gas is probably hydrogen ($H_{2(g)}$)

Answers to Making Connections Questions

35. The density of carbon dioxide is higher than the density of air. The lake was in a volcanic caldera, so the carbon dioxide that bubbled to the surface lay in the volcanic crater since carbon dioxide will sink to the bottom of a sample of air. This displaced the oxygen and asphyxiated the people and livestock.

36. Most of the gases used in a hospital are not toxic, however they must still be handled with caution as they will be stored under pressure. Gases will expand when heated, so canisters must be strong enough to withstand pressure changes that occur with changes in temperature and should be stored in areas where moderate temperatures will be maintained. Since oxygen is flammable, it must not be used around open flame.

37. A low pressure air mass will rise above a high pressure air mass because of the difference in pressure. As the low-pressure air mass rises, it will cool since temperatures are cooler at higher altitude. This will lead to the condensation of the air in the air mass which leads to precipitation.